

Moles, Stoichiometry, and Calculations Review

***BRING YOUR CALCULATOR FOR THE QUIZ!!!

Avogadro's Constant (N_A) and the Mole

- Mole originally defined as number of entities in 12.0 g of Carbon-12
- Now know there are 6.02×10^{23} (N_A) atoms in one mole of any element
- Also, 6.02×10^{23} molecules in one mole of any compound

Molar Mass (M)

Total mass of an element or compound (amu) = molar mass in grams (g)

e.g. What is the mass of one mole of $\text{NaC}_2\text{H}_3\text{O}_2$?

(refer to periodic table to get **average** atomic mass of each element- which is a weighted average of each element's different isotopic masses)

$$\begin{aligned}\text{mass (grams)} &= (1 \times 22.99) + (2 \times 12.01) + (3 \times 1.007) + (2 \times 16.00) \\ &= 82.03 \text{ g}\end{aligned}$$

Calculations involving the Mole concept

1) Converting Mass to Amount in Moles:

$$\text{number of moles (mol)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \quad \text{OR} \quad \boxed{n = \frac{m}{M}}$$

2) Converting Amount in Moles to Mass:

$$\boxed{m = n \times M} \quad (\text{rearrangement of formula from above})$$

Percentage composition

The Law of Definite Proportions: the elements in a compound are always present in the same proportions by mass. (e.g. H_2O is always 11.2% H and 88.8% O by mass)

$$\boxed{\% \text{ of one element} = \frac{\text{mass of that element}}{\text{mass of the sample}} \times 100\%}$$

Masses are found in a lab using a **Mass Spectrometer**:

- samples are ionized, then forced through a magnetic field and an angled tube
- lighter samples divert more from their original path and heavier samples divert less

Empirical versus Molecular Formulas

- a compound's **empirical formula** (simplest formula) is based on % composition
- the **molecular formula** (actual formula) will be a multiple of the empirical formula

e.g. HO is the empirical formula for hydrogen peroxide, but the actual molecular formula is H₂O₂.

Stoichiometry

-Balanced chemical reactions tell us the ratio of MOLECULES that react or the ratio of MOLES of molecules that react.

Limiting and Excess Reagents

If you only have a certain amount of each reactant (your starting materials), one will be the **limiting reagent** and one will be the **excess reagent** (ie. you have leftovers).

Percentage Yield

-Can use balanced chemical equations to calculate the **theoretical yield** of a product
 -When compared to the **actual yield** (what you produce in a lab), you can calculate the **percentage yield**:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Example Short-Answer Question:

NOTE: On the quiz we will provide a table like the one shown below including the balanced chemical equation, formulas, and numbered steps in the upper left corners of each cell.

1. How many grams of NaHCO₃ react exactly with 10.0 g of HC₂H₃O₂? **14.0 g**

Balanced chemical equation	HC ₂ H ₃ O ₂ (aq)	+ NaHCO ₃ (aq) →	Products
¹ Mole ratio	1	to 1	-
Mass (g) $m = n \times M$	² 10.0 g (Given)	⁶ = 0.167 mol x 84.008 g/mol = 14.0 g	-
Molar mass (g/mol) M	³ = (4 x 1.007) + (2 x 12.01) + (2 x 16.00) = 60.048 g/mol	³ = (1 x 22.99) + (1 x 1.007) + (1 x 12.01) + (3 x 16.00) = 84.008 g/mol	-
Moles (mol) $n = \frac{m}{M}$	⁴ = $\frac{10.0 \text{ g}}{60.048 \text{ g/mol}}$ = 0.167 mol	⁵ Mole ratio 1:1 therefore 0.167 mol	-

Step 1: Identify your mole ratio (first row) from the balanced chemical equation.

Step 2: Write in your given mass(es) of reactants (second row).

Step 3: Calculate the molar masses of each substance (third row) using periodic table.

Step 4: Calculate the number of moles of the reactant whose mass is given (fourth row).

Step 5: USE mole ratio from step 1 and the calculated number of moles from step 4 to calculate number of moles of other substance.

Step 6: Calculate the mass of the substance asked for in the question.